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Evaluating the Effect of Temperature on Back Titration Accuracy: A Study of Modern Collision Theory and Calcium Carbonate Content in Eggshells

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ABSTRACT

This investigation examines the effect of temperature on the outcome of a back titration used to determine the percentage composition of calcium carbonate in Gallus gallus domesticus eggshells. While collision theory predicts that increasing temperature enhances reaction rates through more frequent and energetic particle collisions, this study evaluates whether such changes meaningfully influence titration results.

A back titration method was employed in which powdered eggshell samples were reacted with excess hydrochloric acid at three controlled temperatures (14°C, 22°C, and 35°C). The remaining acid was subsequently titrated with standardized sodium hydroxide solution. The volume of titrant required to reach the endpoint was recorded and used to calculate the amount of calcium carbonate present.

Results demonstrated minimal variation in titrant volume across temperature conditions, with a weak positive trend and a negligible coefficient of determination ($R^2 \approx 0.005$). These findings indicate no significant correlation between temperature and titrant volume, suggesting that temperature does not meaningfully affect the measured outcome in this system.

The hypothesis, based on collision theory, was therefore not supported. However, this does not invalidate the theory; rather, it highlights a conceptual limitation, as back titration measures the extent of reaction rather than the reaction rate. Significant experimental errors, including inconsistent endpoint detection and incomplete sample preparation, contributed to a high percent error (97%).

Overall, this study underscores the importance of aligning experimental design with theoretical frameworks when investigating kinetic effects.

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INTRODUCTION

In chemistry, a titration is a quantitative analytical technique used to determine the concentration of an unknown substance through reaction with a standardized solution (Murillo, 2021). Common forms of titration include acid-based, redox, and precipitation titrations. In acid-based titrations, an indicator or pH probe is commonly used to determine the endpoint of neutralization. Redox titrations rely on oxidation-state changes, while precipitation titrations involve the formation of an insoluble product. Titrations are typically conducted in two stages: a scout titration to estimate the endpoint, and a subsequent trial to improve precision (Purdue University, 2019).

In this experiment, operators participated in a back titration experiment in order to determine the percentage amount of calcium carbonate in *Gallus gallus domesticus*. In chemistry, a percentage amount, or percent composition, "tells you by mass what percent of each element is present in a compound" (Belford, 2019). The presence of calcium carbonate in eggshells is biologically important because it provides a rigid protective barrier for the developing embryo; however, calcium carbonate analysis also has broader industrial and agricultural significance. Calcium carbonate content is commonly evaluated in food science, fertilizer production, environmental chemistry, and pharmaceutical applications due to its role in structural stability, nutrient supplementation, and material composition. Additionally, the strength of the calcium carbonate shell serves as a defense mechanism against bacterial contamination. Due to its calcium carbonate content, crushed eggshell powder works as a high-quality fertilizer to increase soil calcium, prevent blossom-end rot in plants, and strengthen plant cell walls. Studies suggest the calcium in eggshells may be up to 64% better absorbed than pure calcium carbonate, partly due to trace minerals and proteins present. To use eggshells efficiently, it is recommended to wash, dry, and grind them into a fine powder, which allows the calcium to be utilized effectively (Bao et al., 1997).

Furthermore, in this experiment, operators dealt with the importance of the chemical reactions present in all the solutions made. For example, the operators attempted to make a dilution by diluting the base present (NaOH) to ensure its strength would not overpower the acid/eggshell solution and cause an unsuccessful and uncontrolled neutralization reaction. The researchers also created a solution between the acid, HCl, the egg shell powder, $CaCO_3$, and water, H_2O, in order to properly produce an analyte. Also, the operators investigated the solution that resulted from the back titration, which included a completed reaction between the titrant and the analyte. The balanced chemical equations for both of these reactions are below:

- Eggshell/Acid Solution: $1CaCO_3(s) + 2HCl(aq) \rightarrow 1H_2O(l) + 1CaCl_2(aq) + 1CO_2(g)$
- Resulting Titration Solution: $1NaOH(aq) + 1HCl(aq) \rightarrow 1H_2O(l) + 1NaCl(aq)$

This experiment deepened the operator's understanding of dilutions, solutions, chemical acid/base reactions, and overall titrations because it challenged operators to deduce the unknown concentration and percent composition of calcium in eggshells. The understanding of energy and its effect on molecular motion is foundational to collision theory and chemical kinetics. According to Collision Theory, proposed by Maxx Trautz in 1916 and William Lewis in 1918, increasing temperature increases average kinetic energy, causing particles to move more rapidly and collide more frequently (Soult, 2023). These changes primarily affect reaction rate rather than the final extent of reaction. Therefore, this investigation examines whether temperature-related kinetic differences meaningfully influence the measured endpoint of a back titration system. Overall, the operators hypothesized that increasing temperature would accelerate molecular collisions and reaction kinetics, potentially influencing the measured titrant volume required to reach the endpoint.

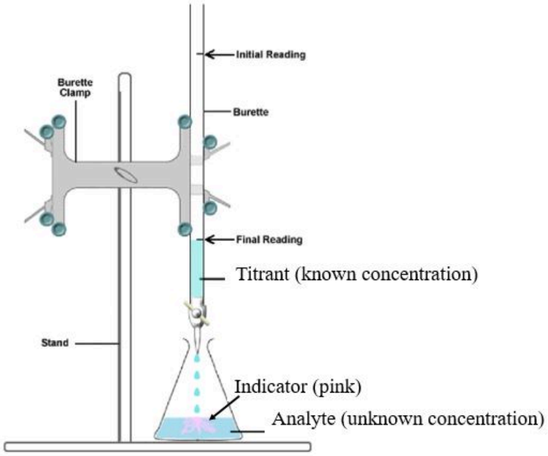

Figure 1: Titration Setup (LibreTexts, 2023).

VARIABLES

Independent and Dependent Variables

The independent variable in this experiment is the temperature of the Hydrochloric Acid (HCl). The temperature is measured in degrees Celsius (°C) and will stay consistent in units for accurate calculations. As a result, the dependent variable is the time taken for complete dissolution/neutralization titration.

Controlled Variables

Variable	Explanation	Method of control
Room environment (temperature, lighting, etc.)	By keeping the room environment constant, researchers ensured that the data recorded was not affected by confounding variables such as differences in light or temperature. If the temperature	Operators turned on and kept the lights in the laboratory on, and maintained the classroom temperature set to 22.2°C.

	varied, the data collected may have changed. Similarly, if the lighting changed, operators may have experienced difficulty in accurately seeing the color of the indicator as the reaction reached its end.	
Mass of eggshell powder	In utilizing a constant amount of eggshell powder, operators are able to ensure the time measured to conduct the neutralization was not affected by the amount of the eggshell, but by the temperature solely.	The operators will obtain 2.250g (±0.001g) of egg shell, uncrushed (broken), and then grind it with a mortar and pestle until a uniform powder. Then, after adding the HCl to the crushed egg shells, the operators will use approximately 12.5 mL (± 0.5)of the eggshell acid solution.
Concentration and Volume of HCl (e.g., 25 mL of 1.0 M)	In keeping consistent measurements of the concentration and volume of HCl, operators are able to ensure the reaction is conducted properly and equally each time, and that the change in the concentration or volume of the HCl is what is causing a change in the reaction speed.	The operators will properly measure 25 mL (±0.5)of 1.0 M HCl for each trial.

Table 1: Controlled variables.

MATERIALS

Item	Quantity and Volume	Uncertainty
Goggles	1 pair	N/A
Apron	1	N/A
Gloves	1 pair	N/A

Eggshells	2.250 g	(±0.001g)
Scale	1	(±0.001g)
Mortar and Pestle	1	N/A
Weigh Boat	2	N/A
Scoopula	1	N/A
Graduated cylinder	1 ; 50mL	(±0.5mL)
Funnel	1	N/A
Beaker	1; 50mL	±5mL
Glass Stirring Rod	1	N/A
Graduated Cylinder	1; 100mL	±12.5mL
Flask	10; 200mL	±10mL
De-Ionized Water	1 squeeze bottle; 100mL	±12.5mL
Phenolphthalein Indicator	1 dropper bottle	N/A
Burette	1; 50mL	±5mL
Ringstand	1	N/A
Burette Clamp	1	N/A
White Piece of Paper	1	N/A
Waste Beaker	1; 50mL	±5mL
Hot Plate	1	N/A
Ice	50 Cubes	N/A
Lab Tray	1; approximately 2 feet by 3 feet	N/A
Thermometers	4	±0.5°C
Hot Vessel Clamps	1	N/A

Table 2: Materials utilized in the experiment.

EXPERIMENTAL PROCEDURE

Pre-Procedures – Day 1

1. All safety equipment, including goggles, gloves, and aprons, was obtained prior to experimentation.
2. All materials and laboratory equipment necessary for the investigation were collected and organized (refer to *Figure 3*).
3. The eggshell solution was prepared according to the following procedure:
    a. Approximately 2.250g (±0.001g) of uncrushed eggshells were obtained
    b. The eggshells were ground using a mortar and pestle until a uniform powder was achieved.
    c. A 2.250g (±0.001g) sample of the crushed eggshell powder was transferred into a 1000ml (±25 mL) flask.
    d. A 1.0 M HCl solution was prepared by diluting concentrated hydrochloric acid according to the following calculation:
        i. (3)(x)=(1.0)(130ml) = 43.3 mL HCl (±0.5mL)
    e. Approximately 43.3 mL (± 0.5 mL) of HCl and 86.7 mL (± 0.5 mL) of deionized water were added to the flask containing the eggshell powder.
    f. The mixture was stirred thoroughly using a glass stirring rod.
    g. The solution was stored in a controlled environment for approximately 96 hours before titration.
4. The NaOH solution was prepared according to the following procedure:
    a. The amount of NaOH required to prepare the solution was calculated using the molarity equation:
        i. M = n/L or n = CV
        ii. 1.0=x/(0.2L) = 0.2 moles → (0.2 moles)(39.998 grams) = 8.000g (±0.001g) NaOH
    b. Approximately 8.000 g (±0.001g) of NaOH was measured.
    c. The NaOH was dissolved in 192 mL (±0.5mL) of deionized water within a 100mL (±25 mL) beaker.
        i. This procedure was repeated to prepare a second NaOH solution.
        ii. Both NaOH solutions were combined in a larger beaker and stirred gently using a glass stirring rod.
        iii. Approximately 30 mL of the combined solution was diluted with 270 mL of deionized water to create the NaOH solution utilized during titration.
5. The NaOH solution was stored in a controlled environment for approximately 96 hours before use.

Procedures – Day 2

1. Appropriate safety equipment, including goggles, gloves, and aprons, was worn throughout the experiment.
2. All required laboratory equipment and materials were assembled before experimentation (refer to *Figure 3*).
3. The titration apparatus was assembled according to the following procedure:

   a. The 50 mL burette was rinsed with deionized water, followed by a small quantity of NaOH solution in order to prime the apparatus.
      i. Deionized water was introduced through the burette while the apparatus was rotated to rinse all internal surfaces.
      ii. A small quantity of NaOH solution was then added and rotated throughout the burette interior.
   b. The burette was secured to a ring stand using a burette clamp
   c. Using a funnel, the burette was filled with NaOH solution to approximately the 0 mL mark.
   d. Air bubbles were removed from the burette tip by releasing small quantities of solution into a waste beaker while gently tapping the nozzle.
   e. The initial burette volume was recorded before titration.
4. Ten flasks containing eggshell acid solution were prepared according to the following procedure:
   a. Approximately 12.5 mL of the eggshell acid solution was transferred into a 200 mL flask
   b. Two or three drops of phenolphthalein indicator were added to the solution, and the flask was gently swirled.
   c. Steps 4(a) and 4(b) were repeated until ten prepared flasks were obtained.
   d. Each flask was labeled appropriately to distinguish room-temperature, cold, hot, and scout titration trials.
5. Temperature-controlled environments for the hot and cold trials were established as follows:
   a. A cold-water bath was prepared using approximately 4 L of water and 50 ice cubes within a laboratory tray.
      i. The temperature of the ice bath was monitored using a thermometer until a stable temperature of 14°C (± 0.5°C) was achieved.
      ii. During cold trials, flasks remained within the ice bath throughout titration to maintain consistent temperature conditions.
   b. A hot plate was used to establish the hot-temperature trials
      i. The hot plate was heated to an intermediate setting before experimentation
      ii. The temperature of the solutions was monitored until approximately 35°C (± 0.5°C) was achieved.
      iii. During hot trials, flasks remained on the hot plate throughout titration to maintain temperature consistency.
6. A piece of white paper was placed below the burette on the ring stand, in order to ensure the color of the solution was easily visible.
7. Room-temperature titrations were conducted using flasks R1-R3.
   a. The initial temperatures of all room-temperature solutions were recorded.
   b. Each flask was positioned beneath the burette on top of a white sheet of paper.
   c. NaOH solution was gradually released into the flask while the solution was continuously swirled to ensure mixing.
   d. The titration was stopped once a faint pink endpoint persisted for approximately 30 seconds.
   e. Final burette volumes were recorded for each trial.

   f. This procedure was repeated for all room-temperature trials.
8. Cold-temperature titrations were conducted using flasks C1-C3.
   a. The initial temperatures of all solutions were recorded.
   b. Each flask was placed in the pre-made ice bath until the proper temperature was achieved.
   c. Each flask was positioned beneath the burette on top of a white sheet of paper.
   d. NaOH solution was gradually released into the flask while the solution was continuously swirled to ensure mixing.
   e. The titration was stopped once a faint pink endpoint persisted for approximately 30 seconds.
9. Hot-temperature titrations were conducted using flasks H1-H3.
   a. The initial temperatures of all solutions were recorded.
   b. Each flask was heated on the hot plate until the proper temperature was achieved.
   c. Each flask was positioned beneath the burette on top of a white sheet of paper.
   d. NaOH solution was gradually released into the flask while the solution was continuously swirled to ensure mixing.
   e. The titration was stopped once a faint pink endpoint persisted for approximately 30 seconds.

<u>Procedures – Clean Up</u>

1. All chemical solutions were disposed of according to Flinn Disposal Method #26b (refer to *Table 3*).
2. Laboratory equipment was rinsed thoroughly using soap and water.
3. The hot plate was unplugged and allowed to cool to room temperature before storage.
4. The ice bath was emptied, and the laboratory tray was dried before storage.
5. All equipment was returned to its designated laboratory location, and work surfaces were cleaned.
6. Safety equipment was removed appropriately, and all disposable materials were discarded properly.

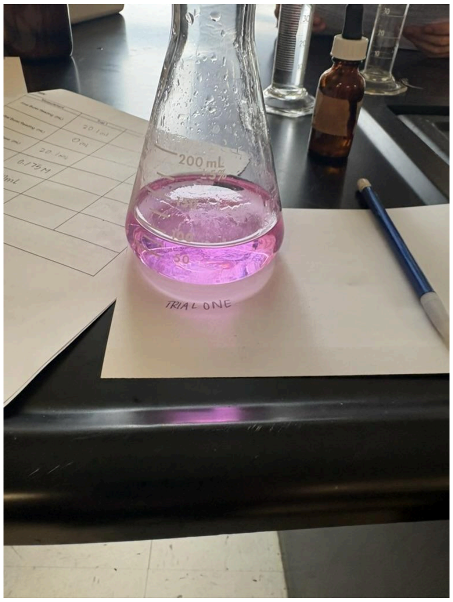

Figure 2: Example of the Color Change Expected for the Scout Titration.

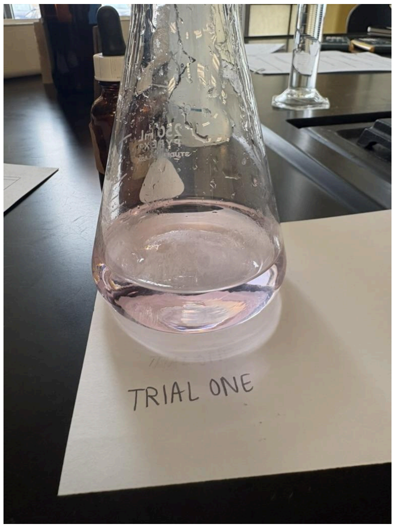

Figure 3: Example of the Color Change Expected for All Trials: 1-9

RISK ASSESSMENT

To ensure the safety of all participants, operators conducted the experiment with proper safety equipment, including goggles and an apron. The goggles were worn at all times, protecting operators from the possibility of glassware shattering due to heat exposure from the hot plate and the risk of splashing chemicals into their eyes. The goggles also provided extra protection against any fumes emitted by the burning of such substances that may have affected vision. Additionally, the aprons ensured that any spilled solutions or chemicals were kept away from the operators' clothes and bodies. There are no ethical concerns with this experiment; however, it is essential to note that flushing large quantities of the chemicals could affect the drainage system and the surrounding aquatic environment.

Moreover, the disposal methods for each elemental mixture are listed below. Please note that all solutions were of small enough quantity to dispose of down the drain safely, but if larger amounts are utilized, refer to *Table 3* for specifications.

Name	Method of Disposal
Sodium Hydroxide (NaOH)	Method #26b: Pour down the drain and flush with appropriate amounts of H_2O.
Hydrochloric Acid (HCl)	Method #26b: Pour down the drain and flush with appropriate amounts of H_2O.
Phenolphthalein	Method #26b: Pour down the drain and flush with appropriate amounts of H_2O.

Table 3: Proper Disposal Methods for Each Compound According to Disposal Methods (2022)

DATA

Qualitative Data

According to Roberts (2023), “Qualitative data is non-numerical data that is produced from qualitative research methods. Qualitative data differs from quantitative data in that the data is descriptive and depicts qualities or characteristics, whereas quantitative data is numerical.” In this experiment, researchers did not observe any qualitative data that is of importance to note.

Essential Equations

ΔV = Change in volume (mL or cm^3)
V_f = Final volume (mL or cm^3)
V_i = Initial volume (mL or cm^3)
$\Delta V = V_f - V_i$
ΔT = Change in Temperature (°C)
T_f = Final temperature (°C)
T_i = Initial temperature (°C)
$\Delta T = T_f - T_i$
n = Moles (mol)
C = Concentration/Molarity ($mol*dm^{-3}$ or M)
V = Volume (mL or L or cm^3 or dm^3)
$n = (C)(V)$

C_1 = The initial concentration of the stock (concentrated) solution
V_1 = The initial volume of the stock solution needed
C_2 = The final concentration of the desired (diluted) solution
V_2 = The final total volume of the diluted solution
Dilution Equation: $C_1V_1 = C_2V_2$
Percent Composition: (part/whole)*100
Percent Error: [(|actual-experimental|)/(actual)]*100
1L = 1000mL

<u>Data Processing/Sample Calculations:</u>

Complete Balanced Chemical Equation for:

- Eggshell/Acid Solution: $1CaCO_3(s) + 2HCl(aq) \rightarrow 1H_2O(l) + 1CaCl_2(aq) + 1CO_2(g)$
- Resulting Titration Solution: $1NaOH(aq) + 1HCl(aq) \rightarrow 1H_2O(l) + 1NaCl(aq)$

Mole Ratio: [(given)(unknown)]/(known)

Creating the HCl solution (1.0M): $(3M)(V_1) = (1.0M)(130mL) \rightarrow V_1 = [(1.0M)(130mL)]/3M = 43.3mL$ $(\pm 0.5mL)$

Calculating the Molar Mass of NaOH:
Na = 1*22.99 = 22.99
O = 1*16.00 = 16.00
H = 1*1.01 =1.01
(1.01)+(16.00)+(22.99) = 39.998 grams → 40 grams

Calculating the amount of NaOH (in grams) needed to create the basic solution: n = CV → n = (1.0M)(0.2L) = 0.2 mol of NaOH → to convert to grams, multiply the moles of NaOH by the molar mass of NaOH → (0.2 moles)(40 grams) = 8.000 grams NaOH

Calculating the change in titrant volume with uncertainties from *Figure 7*: $\Delta V = V_f - V_i \rightarrow 0 \text{ mL} - 2 \text{ mL} = 2mL \rightarrow (\pm 0.05 \text{ mL}) + (\pm 0.05 \text{ mL}) = (\pm 0.1 \text{ mL}) \rightarrow 2mL\ (\pm 0.1 \text{ mL})$

Calculating the average change in volume with uncertainties from *Figure 10*: $(2mL)+(0.2mL)+(0.6mL) = 2.8mL \rightarrow 2.8mL/3 = 0.93mL \rightarrow (\pm 0.5mL)+(\pm 0.5mL)+(\pm 0.5mL) = \pm 1.5mL \rightarrow 0.93mL\ (\pm 1.5mL)$

Calculating the change in temperature with uncertainties from *Figure 11*:
$\Delta T = T_f - T_i \rightarrow 12°C - 14°C = -2°C \rightarrow (\pm 0.5°C) + (\pm 0.5°C) = \pm 1.0°C \rightarrow -2°C\ (\pm 1.0°C)$

Converting milliliters to liters: 12.5mL/1000mL = 0.0125L

Calculating how much HCl actually reacted: First, calculated the amount (in moles) of HCl initially in the flask with the eggshells → HCl: M = 0.1M; V = 12.5mL/1000mL = 0.0125L → n = CV → n = (0.1M)(0.0125L) = 0.00125 mol HCl → Second, calculate the amount (in moles) of HCl in excess through the titration with NaOH. Calculate the amount (in moles) of NaOH needed to reach the endpoint using the average volume (from *Figure 7*) and the molarity of NaOH → M = 0.1; V = 0.93mL/1000mL = 0.00093 L → n = CV → n = (0.1M)(0.00093L) = 0.000093 moles of NaOH → Next, use the balanced equation and stoichiometery (mole ratio) to find the moles of HCl in excess → [(given)(unknown)]/(known) → [(given)(HCl)]/(NaOH) → [(0.000093 mol NaOH)(1 mol HCl)]/(1 mol NaOH) → 0.000093 moles HCl in excess → Now, calculate how much HCl actually reacted by subtracting the initial and excess values of HCl → 0.00125 moles HCl - 0.000093 moles HCl = 0.001157 moles HCl

Calculating the amount (in grams) of calcium carbonate that reacted using stoichiometry: [(given)(unknown)(molar mass)]/[(1)(known)(1mole)] → [(moles HCl)(Coefficient of $CaCO_3$)(Molar Mass of $CaCO_3$)]/[(1)(Coefficient of HCl)(1 mole)] → [(0.001157 mol HCl)(1 $CaCO_3$)(100.09g/mol $CaCO_3$)]/[(1)(2 HCl)(1 mole)] = 0.057902065 grams → 0.0579g $CaCO_3$

Calculating the percentage of calcium carbonate in the eggshells using percent composition: (part/whole)*100 → (grams of $CaCO_3$ that reacted/grams of $CaCO_3$ used)*100 → (0.0579g/2.250g)*100 = 2.57% or 3%

Calculating Percent Error: [(|actual-experimental|)/(actual)]*100 → [(|98-2.57|)/(98)]*100 = 97% error

Quantitative Data

Quantitative data is all numerical data based on equations, measurements, or estimation (*Quantitative Data*, 2021). In this experiment, operators evaluated the change in volume of the titrant utilized, the change in temperature observed throughout the experiment, and the average change in volume and temperature observed throughout the trials.

	Trial 1 (R1)		Trial 2 (R2)		Trial 3 (R3)	
0.1 M NaOH Volume (± 0.05)	Initial (mL)	0	Initial (mL)	0	Initial (mL)	0
	Final (mL)	2	Final (mL)	0.2	Final (mL)	0.6
	Change (mL)	2	Change (mL)	0.2	Change (mL)	0.6
Average Change (mL) (± 1.5 mL)	0.93mL					

Table 4: Raw Data for Change in Volume for the Room Temperature Trials

	Trial 1 (C1)		Trial 2 (C2)		Trial 3 (C3)	
0.1 M NaOH Volume (± 0.05)	Initial (mL)	0	Initial (mL)	0	Initial (mL)	0
	Final (mL)	0.5	Final (mL)	0.4	Final (mL)	0.2
	Change (mL)	0.5	Change (mL)	0.4	Change (mL)	0.2
Average Change	0.37					

(mL) (± 1.5 mL)	

Table 5: Raw Data for Change in Volume for the Cold Temperature Trials

	Trial 1 (H1)		Trial 2 (H2)		Trial 3 (H3)	
0.1 M NaOH Volume	Initial (mL)	0	Initial (mL)	0	Initial (mL)	0
(± 0.05)	Final (mL)	0.4	Final (mL)	0.5	Final (mL)	0.7
	Change (mL)	0.4	Change (mL)	0.5	Change (mL)	0.7
Average Change (mL) (± 1.5 mL)	0.53					

Table 6: Raw Data for Change in Volume for the Hot Temperature Trials

Independent Variable	Trial 1 (R1)		Trial 2 (R2)		Trial 3 (R3)	
Temperature (°C) (±0.5°C)	Initial (°C)	22	Initial (°C)	22	Initial (°C)	22
	Final (°C)	22	Final (°C)	22	Final (°C)	22
	Change (°C)	0	Change (°C)	0	Change (°C)	0
Average Change (°C) (± 1.5°C)	0					

Table 7: Raw Data for Change in Temperature for the Room Temperature Trials

Independent Variable	Trial 1 (C1)		Trial 2 (C2)		Trial 3 (C3)	
Temperature (°C) (±0.5°C)	Initial (°C)	22	Initial (°C)	22	Initial (°C)	22
	Final (°C)	14	Final (°C)	14	Final (°C)	14

	Change (°C)	-8	Change (°C)	-8	Change (°C)	-8
Average Change (°C) (± 1.5°C)	-8					

Table 8: Raw Data for Change in Temperature for the Cold Temperature Trials

Independent Variable	Trial 1 (H1)		Trial 2 (H2)		Trial 3 (H3)	
Temperature (°C) (±0.5°C)	Initial (°C)	22	Initial (°C)	22	Initial (°C)	22
	Final (°C)	35	Final (°C)	35	Final (°C)	35
	Change (°C)	13	Change (°C)	13	Change (°C)	13
Average Change (°C) (± 1.5°C)	13					

Table 9: Raw Data for Change in Temperature for the Hot Temperature Trials

Trial	Initial Temp (°C) (±1.5°C)	Final Temp (°C) (±0.5°C)	Change in Temp (°C) (±0.5°C)	Starting Volume of Burette (mL) (± 0.05)	Ending Volume of Burette (mL) (± 0.05)	Change in Volume (mL) (± 1.5mL)
1R	22	22	0	0	2	2
2R	22	22	0	0	0.2	0.2
3R	22	22	0	0	0.6	0.6
1C	22	14	-8	0	0.5	0.5
2C	22	14	-8	0	0.4	0.4
3C	22	14	-8	0	0.2	0.2
1H	22	35	13	0	0.4	0.4

2H	22	35	13	0	0.5	0.5
3H	22	35	13	0	0.7	0.7

Table 10: Processed Data for Change in Volume and Change in Temperature for All Trials

Trial	Final Temp (°C) (±0.5°C)	Total Titrant Used (mL) (± 0.05)
1R	22	2
2R	22	0.2
3R	22	0.6
1C	14	0.5
2C	14	0.4
3C	14	0.2
1H	35	0.4
2H	35	0.5
3H	35	0.7

Table 11: Data Table for *Figure 4*

ANALYSIS OF DATA

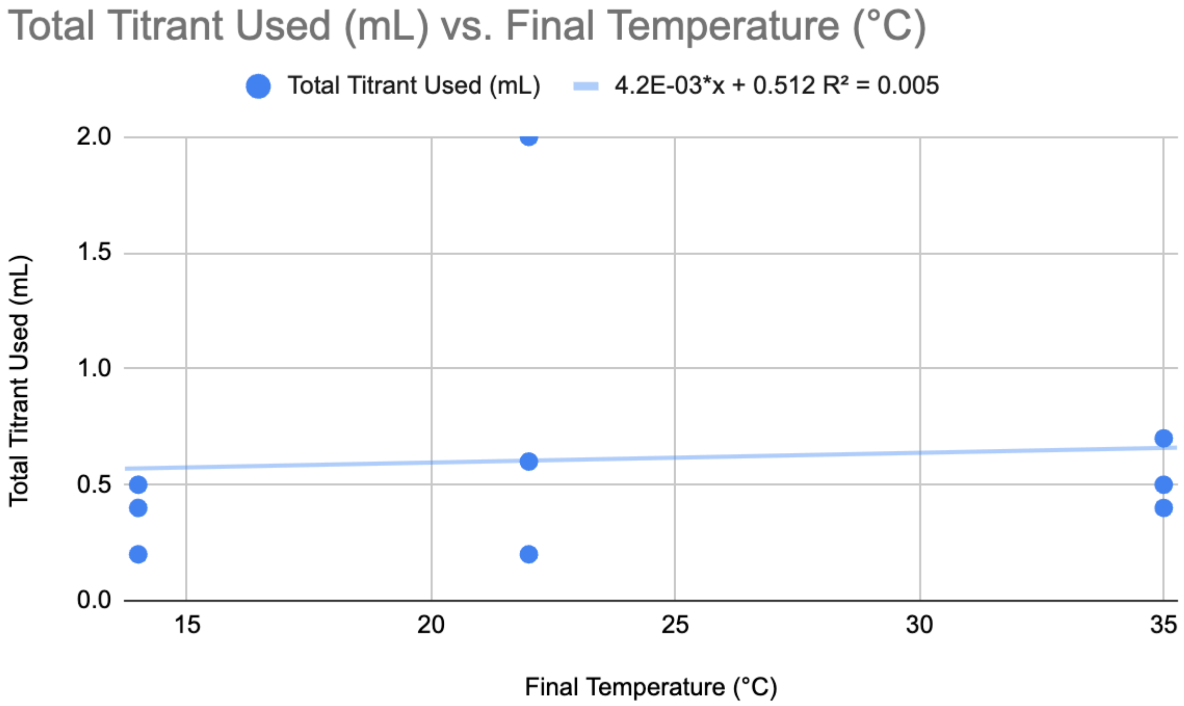

Figure 4: Graph of Total Titrant Used (mL) vs. Final Temperature (°C)

Analytical Discussion of Figure 4

The slope of the line of best fit (0.0042) is very small and positive, meaning that as temperature increases, the amount of titrant used increases very slightly. Therefore, there is almost no meaningful relationship between temperature and the amount of titrant utilized because the slope increases in extremely small increments, the points are widely scattered, and there is no clear upward or downward trend. An R^2 value measures the proportion of variance in the dependent variable, as predicted by the independent variables in the linear regression (Khan Academy, 2017). It is a value between zero and one; the closer the number is to one, the better the fit for the model. The graph in *Figure 15* depicts an R^2 value that essentially means that only 0.5% of the change in titrant used is explained by the temperature, suggesting absolutely no correlation between the two variables. Collision theory says that as temperature increases, particles move faster, collisions happen more often, more collisions have enough energy to react, and the reaction rate increases. If temperature had strongly affected the reaction in this experiment, the graph in *Figure 15* would be expected to demonstrate a clear trend in titrant used with temperature; however, the graph in *Figure 15* shows almost no trend and an extremely weak correlation. The data does not strongly support the predicted temperature-dependent effect proposed by collision theory, but also does not definitively disprove it, because it simply suggests that temperature may not have significantly affected the measured outcome, or experimental error masked the effect of temperature. There is an outlier in this graph that appears to have utilized much more titrant than the other trials used, most likely due to the fact that the operators did not complete the suggested scout titration, but instead used the first room temperature trial as a scout titration, which obviously skews the results. Outliers, especially in this case, matter because they can distort the line of best fit, reduce the accuracy of the model, and lower the R^2. If that specific point were removed, the trend might look slightly different–but given the low R^2, the relationship would likely still be weak. The line of best fit essentially suggests that for every 1°C increase in temperature, the tyrant used increases by only about 0.0042mL, a change that is so small it is practically meaningless.

Error Discussions

Percent Error Calculations:

According to the Journal of the Royal Society, the average amount of calcium carbonate in eggshells is 98% (McClelland et al., 2021).

Calculating Percent Error: [(|actual-experimental|)/(actual)]*100 → [(|98-2.57|)/(98)]*100 = 97% error

Discussion:

The extremely high percent error (97%) suggests that substantial procedural inaccuracies affected the reliability of the results. One major source of error likely originated from the inconsistent eggshell practical size. Because the eggshells were ground manually using a mortar and pestle, larger fragments may not have fully reacted with the HCL solution, reducing the measured amount of reacted $CaCO_3$ and artificially lowering the calculated percent composition. Additionally, the absence of a proper scout titration caused the first room-temperature trial to overshoot the endpoint substantially, producing an outlier that distorted the line of best fit and weakened the R^2 value. Inconsistent endpoint detection also likely contributed to random error because the faint pink phenolphthalein endpoint is subjective and dependent on operator judgment. Furthermore, temperature fluctuations during titration may have reduced experimental consistency, as solutions likely did not remain perfectly at 14°C or 35°C throughout the trials. Together, these procedural limitations increased data scatter, weakened observable trends, and contributed significantly to the high percent error and low correlation observed in the investigation.

CONCLUSION

The operator's hypothesis stated that increasing the temperature of the hydrochloric acid would result in less titrant being required to reach the endpoint because, according to collision theory, particles at higher temperatures move faster, collide more frequently, and are more likely to undergo successful collisions. However, the experimental data do not support this hypothesis. The graph (*Figure 15*) of total titrant used versus temperature produced a very small positive slope and an extremely low R^2 value. The R^2 value of 0.005 indicates that only 0.5% of the variation in titrant volume can be explained by temperature. This demonstrates that temperature had essentially no measurable effect on the volume of NaOH required to neutralize excess HCl. Additionally, the average titrant volumes were: room temperature (22°C), 0.93mL (heavily skewed by an outlier), cold (14°C), 0.37mL, and hot (35°C), 0.53mL. While there is a slight numerical increase between the cold and hot trials, the variation is inconsistent and does not show a clear proportional trend; therefore, the hypothesis is rejected based on experimental evidence. However, rejecting the hypothesis does not invalidate Collision Theory. Instead, it suggests that temperature may not significantly affect the extent of reaction in this system; temperature primarily affects reaction kinetics rather than final reaction extent, and experimental limitations likely masked any temperature-dependent effects.

According to Flowers et al. (2018), one major conceptual limitation is that Collision Theory explains reaction rate, not the final amount of product formed. In this experiment, the back titration measures how much HCl remains after reaction with $CaCO_3$. Given enough time, the reaction would go to completion regardless of temperature. Temperature increases reaction speed, but not necessarily the total moles reacted once equilibrium is reached.

The eggshell/HCl mixture was left for 96 hours; all reactions between the two likely reached completion before the titration began. Thus, the temperature during titration would not significantly alter the total HCl remaining, suggesting that the experimental design may not have been appropriately aligned with the theory being tested. One major limitation of this investigation was the fact that only three temperature values were tested, and the sample size within those temperatures was very small (only three trials each). Furthermore, back titrations measure remaining acid, not reaction speed directly, so the operators were challenged to find that independently. An extremely high percent error (97%) suggests major procedural inaccuracy; because the volumes measured were very small, even tiny reading errors represent a large percentage of the measurement.

Finally, to improve and extend the experiment, operators could instead measure the reaction rate instead of the final amount in order to directly test collision theory. By testing 5°C intervals from 5°C to 45°C, operators could increase the accuracy of the regression. Using a physical pH probe or a physical indicator instead of phenolphthalein could also increase precision. Utilizing a mechanical grinder and a sieve to ensure uniform particle size among the eggshells would tremendously improve the eggshell preparation, and remove availability for error. Lastly, conducting additional trials per temperature condition could reduce the impact of random error.

In the end, the hypothesis was rejected because the data demonstrate no significant correlation between temperature and titrant volume, as evidenced by the R^2 value. However, this rejection does not invalidate Collision Theory, but rather highlights a misalignment between theoretical expectations and experimental design. Additionally, substantial systematic errors likely contributed to the extremely high percent error of 97%. Therefore, while the experimental results do not support the hypothesis, they are more reflective of the limitations of titrations than the inaccuracy of Collision Theory itself.

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